Chemistry B – Final Exam Review Packet   Spring 2017

The final exam will count as approximately 15% of your final grade in Chemistry B.

Exam Format:

- Multiple choice ~35 questions
- Free Response/Calculations: ~35 points
- The exam will cover material from the entire trimester, but will emphasize material from the last unit including molecular structure, intermolecular forces, and heating curves.

Materials you need to bring:

Calculator, #2 pencil, your Periodic Table with references on the back.

Materials provided:

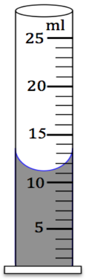
VSEPR geometry sheet, any constants or values not provided on the periodic table, scantron, and scratch paper.

Topics covered on the exam, and skills that may be assessed:

☼ **Analyze and Report Measurements using Significant Figures and Units.**	**Chapters 3 & 4**
	Accurately report measurements using the correct number of significant figures and units.
	Report calculated values using the appropriate units and significant figures.
☼ **Molar Conversions, Percent Composition, Empirical and Molecular Formulas**	**Ch 7: Pg. 170 – 197**
	Use the Periodic Table to determine the molar mass of an element or compound.
	Determine the percent composition by mass of each element in a compound.
	Determine the empirical formula of a compound given the percent composition.
	Determine the molecular formula given molecular mass and an empirical formula.
☼ **Kinetic Theory and Gas Laws**	**Ch 12: Pg. 327 – 355**
	Describe the motion of gas particles and interpret observable changes in temperature, pressure, and volume.
	Calculate the resulting temperature, pressure, volume, or number of moles of gas using gas law equations.
☼ **Molarity and Solutions**	**Ch 18: Pg. 509 – 515**
	Complete calculations involving molarity of solutions. Apply the equations for molarity and dilutions.
☼ **Stoichiometry**	**Ch 9: Pg. 238 – 259**
	Interpret a balanced chemical equation and use it to calculate how much product will be formed (in particles, mass, volume, or moles) when given the amount of another substance in the reaction.
	Identify which reactant is limiting and which reactant is in excess as well as calculate how much product can be produced (theoretical yield) when given quantities of reactants.
	Use a theoretical yield and actual yield to calculate the percent yield of a chemical reaction.
☼ **Covalent Compounds**	**Ch 16: Pp. 436 – 459**
	Draw the electron dot structure for compounds with a single central atom.
	Use VSEPR theory to identify the geometric shape of a molecule.
	Use electronegativity and molecular symmetry to determine whether bonds or compounds are polar or nonpolar.
☼ **Intermolecular Forces and Aqueous Solutions.**	**Chapter 16 & 17 (pp. 460 – 466 & 474 – 477)**
	Predict the dominant intermolecular forces in a given molecule (dispersion forces, dipole interactions, and/or hydrogen bonds) based on its structure and the presence of polar bonds.
	Discuss how molecular structure and intermolecular attractions determine observable properties including solubility, adhesion, cohesion, surface tension, viscosity, volatility, melting point, and boiling point.
☼ **Thermochemistry**	**Ch 11: Pg. 293 – 218**
	Label and interpret heating and cooling curves for different materials.
	Describe the direction of heat flow during different chemical and physical processes.
	Identify chemical or physical processes as exothermic (releasing heat) or endothermic (absorbing heat).
	Identify the phases and changes in phase on a heating curve and calculate changes in heat associated with changes in temperature and phase.

Unit 1: Scientific Measurement & Chemical Quantities

Measurement & Significant Figures: Ch 3-4



1. Record the volume of liquid pictured to the left. Use the correct significant figures and units.
2. Someone else measures out 32.3 mL of liquid and adds it to the liquid you measured in problem 1, above. Calculate the total volume of the combined solution and record the value using significant figures and units.

Molar Conversions, Percent Composition, Empirical and Molecular Formulas: Ch 7 (pg. 170 – 196)

1. Determine the number of representative particles in each of the following:

a. 1.00 mol $Al(OH)_3$		c. 1.00 mol Hf	
b. 1.00 mol $Ca(C_2H_3O_2)_2$		d. 1.00 mol $C_6H_{12}O_6$	

2. Determine the number of moles of each of the following:

a. 6.022×10^{23} $Al(OH)_3$ particles		c. 178.5 g of Hf	
b. 22.4 L of CO_2 (@STP)		d. 180.156 g of $C_6H_{12}O_6$	

3. Find the empirical formulas for the given molecular formulas. The first one has been done as an example.

a. C_8H_{18}	÷ ***2*** C_4H_9	c. $C_2H_4O_2$		e. C_6H_5N	
b. N_2H_4		d. P_4O_{10}		f. Se_3O_9	

4. Determine the percent composition by mass of each element in the following compounds:

   a. LiCl

   b. $Al(NO_3)_3$

   c. $Hg(OH)_2$

5. Use percent composition by mass to determine the empirical formula of each of the following compounds:

   a. A compound that is 34.43% iron and 65.57% chlorine.

   b. A compound that contains 85.6% carbon and 14.4% hydrogen.

   c. A compound that is 45.9% potassium, 16.5% nitrogen, and 37.6% oxygen.

6. Determine the molecular formulas for each of the following:
   a. A compound with a molecular mass of 78.1 g/mol and an empirical formula of CH

   b. A compound with a molecular mass of 32.1 g/mol and an empirical formula of NH_2

   c. A compound with a molecular mass of 88.0 g/mol and an empirical formula of C_2H_4O

Unit 2: The Behavior of Gases – Ch 12 (pp. 327 – 355)

1. Draw a graph showing the general trend for each of the following gas law relationships and identify the whether the relationship is direct or inverse.

$$P_1V_1 = P_2V_2$$

Direct or **Inverse**

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Direct or **Inverse**

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Direct or **Inverse**

2. A rigid container holds a gas at a pressure of 55 kPa and a temperature of –100.0°C. What will the pressure be when the temperature is increased to 200.0°C?

3. A helium balloon has a volume of 25.0 L at 102.0 kPa and 24 °C. Determine its volume at standard temperature and standard pressure(STP).

4. Calculate the grams of oxygen (O_2) in a 12.5 L tank if the pressure is 25,325 kPa and the temperature is 22.0 °C.

Unit 3, Part 1: Molarity and Solutions – Ch 18 (pp. 509 – 515)

1. Determine the molarity of a 100. mL solution made by dissolving 4.95 g NaCl in water.

2. Determine the mass in grams of H_2SO_4 in 15 mL of a 2.4 M H_2SO_4 solution.

3. What volume of 12 M HCl solution will contain 1.0 moles of HCl?

4. Determine the final concentration of a solution made by diluting 23.4 mL of 6.0 M NaCl stock solution to a final volume of 250. mL

Unit 3, Part 2: Stoichiometry – Ch 9 (pp. 238 – 259)

1. Balance the chemical equation below, and use it for the questions 2 through 6:

$$\_\_\_\_C_2H_6 + \_\_\_\_O_2 \rightarrow \_\_\_\_CO_2 + \_\_\_\_H_2O$$

2. Determine the molar masses (with units) of each reactant and product:

C_2H_6:	O_2:	CO_2:	H_2O:

3. How many moles of CO_2 are formed when 3.7 moles of C_2H_6 are reacted with excess oxygen?

4. Determine the mass of water produced if 64.8 grams of C_2H_6 combust with excess oxygen.

5. How many liters of oxygen are needed to react with 12.5 L of C_2H_6? Assume standard temperature and pressure.

6. What mass of carbon dioxide gas will be produced when 15.6 g of C_2H_6 is reacted with excess oxygen?

   If this reaction were carried out and only 40.6 g of carbon dioxide were produced, what would be the percent yield?

7. Balance chemical equation for the single-replacement reaction between aluminum and iron (II) sulfate, and use it to complete the following problems:

$$\_\_\_\_Al\ (s) + \_\_\_\_FeSO_4 \rightarrow \_\_\_\_Fe + \_\_\_\_Al_2(SO_4)_3$$

8. Determine the molar masses of each reactant and product:

Al :	_______:	____:	__________:

9. Calculate the number of aluminum atoms need to react with 2.56 moles of iron (II) sulfate.

10. How many grams of iron can be produced if 1.25 g of aluminum and 9.05 g of iron (II) sulfate are reacted?

    Which reactant is the limiting reactant? ________________ Which is the excess reactant? ________________
    Determine the grams of unreacted excess reactant that remain after the reaction is complete.

11. In the lab, 0.55 grams of aluminum are reacted with excess iron (II) sulfate. Calculate the percent yield if the reaction produces 1.52 grams of iron.

12. Solid carbon and liquid water react to produce carbon tetrahydride gas and carbon dioxide gas. The balanced chemical reaction is written below.

$$2\ C(s) + 2\ H_2O \rightarrow CH_4(g) + CO_2(g)$$

a. 35.0 g of solid carbon react with excess water. Determine the theoretical yield (in liters) of carbon tetrahydride gas produced at STP.

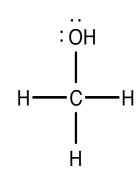
b. How many grams of carbon dioxide can be expected from the reaction if the percent yield is 85.0 %?

Unit 4: Covalent Compounds and Intermolecular Forces – Ch 16 &17 (pp. 436 – 466 & 474 – 477)

1. According to the octet rule, most atoms become more stable when they have ______ valence electrons. The exception to this rule is ____________________, which is most stable with ______ valence electrons.
2. How do you know whether a molecule will experience:
   a. dispersion forces
   b. dipole-dipole attractions
   c. hydrogen bonding
3. State whether the following compounds contain polar covalent bonds, non-polar covalent bonds, or ionic bonds, based on their electronegativities.
   a. KF
   b. SO_2
   c. NO_2
   d. Cl_2
   e. Na_2O
   f. O_2

ΔEN	bond type
0.0 – 0.4	nonpolar covalent
0.4 – 1.0	moderately polar covalent
1.0 – 2.0	very polar covalent
≥ 2.0	ionic

4. Draw the Lewis dot structures for the following compounds, and identify the strongest type of cohesive intermolecular attraction each molecule will experience.
   a. Br_2
   b. CBr_4
   c. CH_2Br_2
   d. CH_3OH



5. Which of the compounds in problem 4 do you expect to have the *highest* boiling point?
6. Predict the order these compounds will evaporate in at room temperature. Which will be the most *volatile?*

   fastest ______________ ______________ ______________ ______________ slowest

7. Define the following terms and explain how they are related to intermolecular attractions.

Cohesion:

Adhesion:

Surface Tension:

8. The figure below indicates the shape of a droplet with high surface tension.
   a. For the droplet pictured, which is stronger, the adhesive forces or the cohesive forces? ____________________
   b. Sketch how the shape of the droplet will change if something is added to weaken the cohesive forces.
   c. Sketch how the shape of the droplet will change if it is put on a surface with stronger adhesive forces.

a. b. c.

Complete the Table: *If a compound has resonance, be sure to draw all possible structures.

Draw the Dot Structure	Draw the 3-D structure	Name the VSEPR Shape, and indicate polarity	*Check*(✓) all forces present & *Circle or box* the ☑ to identify the **strongest** force.	
HF:	3-D Structure:	Shape Name:	dispersion	
			dipole-dipole	
		Polar or Nonpolar?	hydrogen bonding	
PF_3:	3-D Structure:	Shape Name:	dispersion	
		Polar or Nonpolar?	dipole-dipole	
			hydrogen bonding	
SO_2:	3-D Structure:	Shape Name:	dispersion	
		Polar or Nonpolar?	dipole-dipole	
			hydrogen bonding	
XeF_4:	3-D Structure:	Shape Name:	dispersion	
		Polar or Nonpolar?	dipole-dipole	
			hydrogen bonding	
NH_3:	3-D Structure:	Shape Name:	dispersion	
		Polar or Nonpolar?	dipole-dipole	
			hydrogen bonding	
PF_5 :	3-D Structure:	Shape Name:	dispersion	
		Polar or Nonpolar?	dipole-dipole	
			hydrogen bonding	

H_2O :	3-D Structure:	Shape Name: Polar or Nonpolar?	dispersion	
			dipole-dipole	
			hydrogen bonding	
SF_4:	3-D Structure:	Shape Name: Polar or Nonpolar?	dispersion	
			dipole-dipole	
			hydrogen bonding	
PO_4^{3-} :	3-D Structure:	Shape Name: Polar or Nonpolar?	dispersion	
			dipole-dipole	
			hydrogen bonding	
NO_3^{1-} :	3-D Structure:	Shape Name: Polar or Nonpolar?	dispersion	
			dipole-dipole	
			hydrogen bonding	
I_3^{1-} :	3-D Structure:	Shape Name: Polar or Nonpolar?	dispersion	
			dipole-dipole	
			hydrogen bonding	

9. Identify the types of intermolecular forces each of these compounds will exert. Then identify the compounds from the table above that the compound is likely to adhere strongly to.
    a. CH_4
    b. H_2O

Unit 5: Thermodynamics: Ch 11 (pg. 293 – 218)

1. In which direction does heat flow when two objects of different temperatures come into contact with one another? Give an example from your own experience.

2. Is freezing a popsicle an endothermic or an exothermic process? Explain your answer.

3. Complete the following table: Fill in what you'd expect to see for exothermic versus endothermic systems.

	Exothermic	Endothermic
Sign of ΔH_{system}		
Heat flow (in/out of system)		
Measured ΔT of Surroundings		
2 Examples		

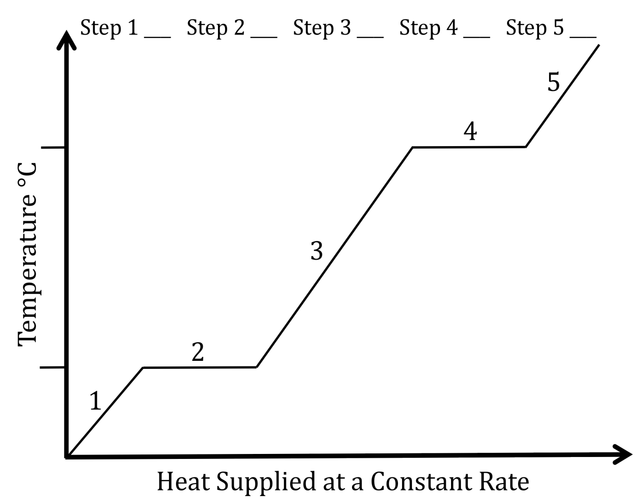
A heating curve is shown to the right.

4. Label each section of the curve with the corresponding phases (s, l, g, etc).

5. Match each step on the heating curve for water to the corresponding behavior (write the letters by the steps).

Description of Behavior

A. Energy is used to increase the temperature of solid ice.
B. Energy is used to increase the temperature of liquid water.
C. Energy is used to increase the temperature of gaseous water (steam).
D. Energy is used to melt ice (S → L).
E. Energy is used to vaporize water (L → G)



6. Identify the steps (1 to 5) on the heating curve above that correspond to each of the terms listed below - some terms refer to multiple steps.
   a. heat of fusion step (s)________________
   b. heat of vaporization step (s)________________
   c. heat of solidification step (s)________________
   d. heat of condensation step (s)________________
   e. latent heat step (s)________________
   f. sensible heat step (s)________________

For the following questions, refer to the table of specific heat values to the right.

Specific Heat $\frac{J}{g \cdot °C}$	
Ethanol	2.44
Mercury	0.14
Hydrogen	14.30
Radon	0.094
Water	4.18

7. Compare the specific heats of ethanol and mercury. Which substance requires less energy to heat to a higher temperature? Why? (Assume equal masses.)

8. Which requires more energy to increase the temperature by 1 °C? Explain why.
   1 g ethanol 1000 g ethanol 1 g mercury 1000 g mercury

9. The element hydrogen has the highest specific heat of all elements. Determine the amount of energy needed to raise the temperature of a 340.0 g sample of hydrogen by 30 °C.

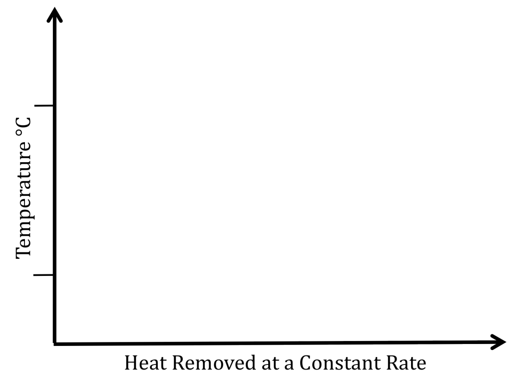
10. Brass is an alloy made from copper and zinc. A 590.0 g brass candlestick has an initial temperature of 98.0°C. When $2.11 x 10^4$ J of energy is removed from the candlestick, its temperature decreases to 6.8 °C. Determine the specific heat of brass.

11. The element radon has the lowest specific heat of all naturally occurring elements. Calculate the change in heat needed to cool 35.0 g of radon by 10.0 °C.

Thermodynamic Properties of Various Substances

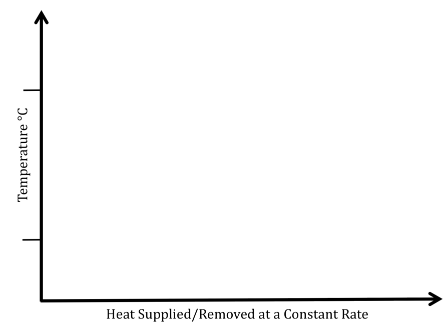
Substance	C_{solid} $\left(\frac{J}{g\cdot°C}\right)$	Melting Point (°C)	ΔH_{fus} $\frac{J}{g}$	$C_{liq,}$ $\left(\frac{J}{g\cdot°C}\right)$	Boiling Point (°C)	ΔH_{vap} $\frac{J}{g}$	$C_{gas,}$ $\left(\frac{J}{g\cdot°C}\right)$
Water	2.10	0.00	334	4.18	100.0	2260	2.00
Ethanol	2.47	– 117	109	2.49	78	838	1.74
Benzene	1.51	5.5	444	1.73	80.1	390	1.06

12. Sketch a ***cooling curve*** for benzene.
    a. Label the temperature axis with the melting point and the boiling point, and identify the phases for each section of the curve.
    b. Using only variables, write the equation you would use to calculate the change in energy when 35.0 g of benzene is cooled from 85.4 °C to 10.2 °C (don't solve them).

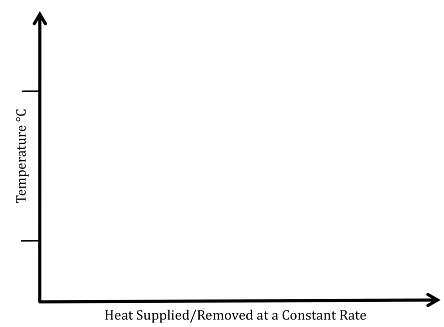


Determine the amount of heat gained or lost during each of the following changes. (Use the values provided above.)

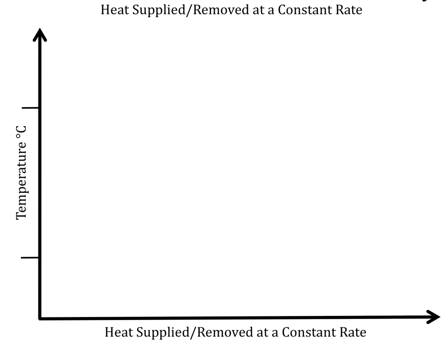
13. 45.0 g of liquid water at 20 °C is converted to steam at 115 °C.



14. 220.0 g of solid water at -35.0 °C is heated to form liquid water at 50.0 °C.



15. 20.0 g of **benzene** at –45.0 °C is heated to 10.5 °C.



16. 5.00 g of **ethanol** at 155 °C is cooled to 60.0 °C.